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Chapter 8
Atoms and Periodic Properties

Slide 2

The origins of quantum physics

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Atomic structure discovered

Ancient Greeks
- Democritus (460-362 BC) - indivisible particles called “atoms”
- Prevailing argument (Plato and Aristotle) - matter is continuously and infinitely divisible

John Dalton (early 1800’s) - reintroduced atomic theory to explain chemical reactions

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Dalton's atomic theory

1. All matter = indivisible atoms
2. An element is made up of identical atoms
3. Different elements have atoms with different masses
4. Chemical compounds are made of atoms in specific integer ratios
5. Atoms are neither created nor destroyed in chemical reactions

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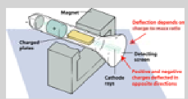
Discovery of the electron

J. J. Thomson (late 1800's)
- Performed cathode ray experiments
- Discovered negatively charged electron
- Measured electron's charge-to-mass ratio
- Identified electron as a fundamental particle



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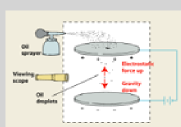
Electron charge and mass

Robert Millikan (~1906)
- Studied charged oil droplets in an electric field
- Charge on droplets = multiples of electron charge
- Charge + Thomson's result gave electron mass



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Early models of the atom

- Dalton - atoms indivisible
- Thomson and Millikan experiments
  - Electron mass very small, no measurable volume
  - What is the nature of an atom's positive charge?
- Thomson's "Plum pudding" model
  - Electrons embedded in blob of positively charged matter like "raisins in plum pudding"

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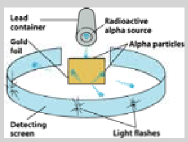
The nucleus



Ernest Rutherford (1907)

- Scattered alpha particles off gold foil
- Most passed through without significant deflection
- A few scattered at large angles
- Conclusion: an atom's positive charge resides in a small, massive nucleus
- Later: positive charges = protons
- James Chadwick (1932): also neutral neutrons in the nucleus

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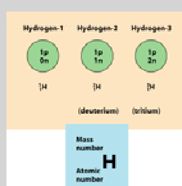
The nuclear atom

- Atomic number
  - Number of protons in nucleus
  - Elements distinguished by atomic number
  - 113 elements identified
  - Number of protons = number of electrons *in neutral atoms*
- Isotopes
  - Same number of protons; different number of neutrons



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Classical “atoms”

Predictions of classical theory

- Electrons orbit the nucleus
- Curved path = acceleration
- Accelerated charges radiate
- Electrons lose energy and spiral into nucleus
- Atoms cannot exist!

Experiment - atoms do exist

$\Rightarrow$ New theory needed

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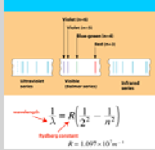
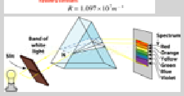
Atomic spectra





Blackbody radiation

- Continuous radiation distribution
- Depends on temperature of radiating object
- Characteristic of solids, liquids and dense gases

Line spectrum

- Emission at characteristic frequencies
- Diffuse matter: incandescent gases
- Illustration: Balmer series of hydrogen lines

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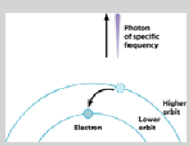
Bohr’s theory

Three rules:

1. Electrons only exist in certain allowed orbits
2. Within an orbit, the electron does not radiate
3. Radiation is emitted or absorbed when changing orbits



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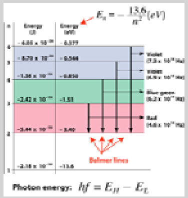
Quantum theory of the atom



- Lowest energy state = "ground state"
- Higher states = "excited states"
- Photon energy equals difference in state energies
- Hydrogen atom example
  - Energy levels
  - Line spectra

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Quantum mechanics

- Bohr theory only modeled the line spectrum of H
- Further experiments established wave-particle duality of light and matter
  - Young's two slit experiment produced interference patterns for both photons and electrons

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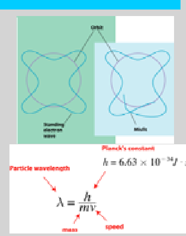
Matter waves

Louis de Broglie (1923)

- Postulated matter waves
- Wavelength related to momentum
- Matter waves in atoms are standing waves



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Wave mechanics

- Developed by Erwin Schrodinger
- Treats atoms as three dimensional systems of waves
- Contains successful ideas of Bohr model and much more
- Describes hydrogen atom and many electron atoms
- Forms our fundamental understanding of chemistry

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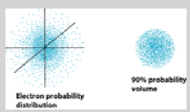
The quantum mechanics model

- Quantum numbers specify electronic quantum states
- Visualization - wave functions and probability distributions
- Electrons delocalized



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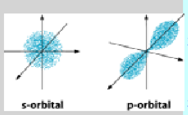
Electronic quantum numbers in atoms



1. Principle quantum number, n
   - Energy level
   - Average distance from nucleus
2. Angular momentum quantum number, l
   - Spatial distribution
   - Labeled s, p, d, f, g, h, …
3. Magnetic quantum number
   - Spatial orientation of orbit
4. Spin quantum number
   - Electron spin orientation

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Electron configuration

- Arrangement of electrons into atomic orbitals
- Principle, angular momentum and magnetic quantum numbers specify an orbital
- Specifies atom's quantum state

- Pauli exclusion principle
  - Each electron has unique quantum numbers
  - Maximum of two electrons per orbital (electron spin up/down)
- Chemical properties determined by electronic structure

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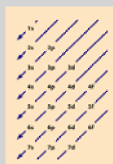
Writing electron configurations

- Electrons fill available orbitals in order of increasing energy
- Shell capacities
  - s = 2
  - p = 6
  - d = 10
  - f = 14
- Example: strontium (38 electrons)



$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$